Structure 1: Models of the Particulate Nature of Matter

IB SL Study Guide

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IB TIP

Data booklet: You can use the IB Chemistry Data Booklet in the exam — all constants, the periodic table, and key equations are provided.

IB Chemistry SL — Atomic Structure & Periodicity

Complete Study Guide

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1. Models of the Particulate Nature of Matter (S1.1)

Chemistry begins with the idea that all matter is made of tiny, discrete particles. This section traces how our model of the atom has evolved from Dalton's simple indivisible sphere to the quantum model used today. Understanding *why* each model changed is just as important as knowing the final picture — IB exams frequently ask you to evaluate the experimental evidence that drove each revision.

Dalton's Atomic Model

John Dalton (1803) proposed the first modern atomic theory:

- All matter is made of tiny, indivisible particles called **atoms**.
- All atoms of the same element are identical in mass and properties.
- Atoms of different elements have different masses and properties.
- Atoms combine in whole-number ratios to form compounds.
- In a chemical reaction, atoms are rearranged — they are never created or destroyed.

Dalton's model explained the law of conservation of mass and the law of definite proportions. However, it could not explain the existence of subatomic particles or why elements emit specific colours of light.

Thomson's Plum Pudding Model

J.J. Thomson (1897) discovered the **electron** through cathode ray tube experiments. He found that:

- Cathode rays were deflected by electric and magnetic fields, proving they were charged particles.
- The charge-to-mass ratio was the same regardless of the metal used for the cathode.

Thomson proposed a model in which the atom was a **sphere of positive charge** with electrons embedded throughout — like plums in a pudding. This was the first model to include subatomic particles, but it was soon overturned by Rutherford's experiment.

Rutherford's Gold-Foil Experiment (Nuclear Model)

Ernest Rutherford (1909–1911) fired a beam of positively charged **alpha particles** (α-particles, helium nuclei) at a very thin gold foil and observed the deflections using a detector screen.

Observations and conclusions:

Observation	Conclusion
Most α-particles passed straight through	Most of the atom is **empty space**
A small fraction were deflected at large angles	There is a small, **dense, positively charged** region — the **nucleus**
A tiny fraction bounced almost straight back	The nucleus is very small but contains most of the atom's **mass**

This experiment **disproved** Thomson's plum-pudding model (which predicted only tiny deflections from a diffuse positive charge). Rutherford's nuclear model introduced a tiny, dense, positive nucleus surrounded by mostly empty space with electrons orbiting at large distances.

> ⚠ **EXAM ALERT**
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> **Exam trap — Rutherford's experiment:** The exam often asks you to explain *which specific observation* led to *which specific conclusion*. "Most particles passed through" proves empty space. "Some deflected at large angles / bounced back" proves a small dense positive nucleus. Do not just say "the nucleus exists" — link the observation to the conclusion explicitly.

The Bohr Model

Niels Bohr (1913) modified Rutherford's model to explain **atomic emission spectra** (see Section 3). Bohr proposed:

- Electrons orbit the nucleus in fixed **energy levels** (also called shells or principal quantum levels), labelled $n = 1, 2, 3, \ldots$
- An electron in a specific energy level has a fixed energy — it does not radiate energy while in that orbit.
- When an electron **absorbs** energy, it jumps to a **higher** energy level (excited state).
- When an electron **falls back** to a lower energy level, it emits energy as a **photon** of light with a specific frequency.

The Bohr model successfully explained the line spectrum of hydrogen. However, it failed to accurately predict spectra of multi-electron atoms, and it could not explain why energy levels were stable.

The Quantum/Wave Model (Modern Atomic Model)

The modern model, developed by Schrödinger, Heisenberg, and others, treats electrons as **waves** rather than particles on defined orbits. Key ideas:

- It is impossible to know both the exact position and momentum of an electron simultaneously (**Heisenberg uncertainty principle**).
- Electrons occupy **orbitals** — regions of space where there is a high probability of finding an electron.
- Orbitals are organised into **sublevels** (s, p, d, f) within each principal energy level.

For IB SL, you need to know the electron configuration model (shells and sublevels) — the full mathematical wave treatment is beyond the syllabus.

Key Definitions — Atoms, Elements, Compounds, Mixtures

Term	Definition
Atom	The smallest unit of an element that retains the chemical properties of that element
Element	A pure substance made of only one type of atom; cannot be broken down by chemical means
Compound	A pure substance made of two or more elements chemically bonded in fixed ratios
Mixture	Two or more substances physically combined; components retain their own properties and can be separated by physical means
Ion	An atom or group of atoms that has gained or lost one or more electrons, giving it an overall charge

2. The Nuclear Atom — Subatomic Particles, Isotopes & Mass Spectrometry (S1.2)

Subatomic Particles

The nucleus contains **protons** and **neutrons**; **electrons** orbit outside the nucleus.

Particle	Symbol	Relative Charge	Relative Mass	Absolute Charge (C)	Absolute Mass (kg)	Location
Proton	p^+	$+1$	1	$+1.60 \times 10^{-19}$	1.673×10^{-27}	Nucleus
Neutron	n^0	0	1	0	1.675×10^{-27}	Nucleus
Electron	e^-	-1	$\frac{1}{1836} \approx 0$	-1.60×10^{-19}	9.109×10^{-31}	Shells (outside nucleus)

MEMORISE THIS

Key nuclear quantities:

- **Atomic number** Z = number of protons (defines the element)
- **Mass number** A = number of protons + neutrons
- **Number of neutrons** = $A - Z$
- In a neutral atom: number of electrons = number of protons = Z
- Notation: ${}^{A}_{Z}\text{X}$ — e.g. ${}^{23}_{11}\text{Na}$ means sodium with $A = 23$, $Z = 11$, so 11 protons, 12 neutrons, 11 electrons

Isotopes

Isotopes are atoms of the same element that have the same atomic number (Z) but different mass numbers (A) — i.e., the same number of protons but different numbers of neutrons.

Examples:

- Carbon: ${}^{12}_{6}\text{C}$ (6 protons, 6 neutrons) and ${}^{14}_{6}\text{C}$ (6 protons, 8 neutrons)
- Chlorine: ${}^{35}_{17}\text{Cl}$ (17 protons, 18 neutrons) and ${}^{37}_{17}\text{Cl}$ (17 protons, 20 neutrons)
- Hydrogen: ${}^{1}_{1}\text{H}$ (protium), ${}^{2}_{1}\text{H}$ (deuterium), ${}^{3}_{1}\text{H}$ (tritium)

Isotopes have identical chemical properties because they have the same number of electrons (and therefore the same electron configuration). They differ in physical properties (e.g. density, rate of diffusion) due to the difference in mass.

⚠ EXAM ALERT

Exam trap — isotope properties: Students often confuse “same chemical properties” with “same physical properties”. Isotopes have the **same** chemical properties (same number of electrons / same electron configuration) but **different** physical properties (different masses). Do not state that isotopes are chemically different — they are not.

Relative Atomic Mass (A_r)

Because most elements exist as a mixture of isotopes, the **relative atomic mass** (A_r) is the **weighted average mass** of all naturally occurring isotopes of an element, relative to $\frac{1}{12}$ of the mass of one atom of carbon-12.

$$A_r = \sum (\text{fractional abundance} \times \text{mass number of isotope})$$

WORKED EXAMPLE

Worked Example 1 — Calculating A_r from isotope data

Chlorine has two naturally occurring isotopes:

- ^{35}Cl: abundance 75.77%
- ^{37}Cl: abundance 24.23%

$$A_r(\text{Cl}) = (0.7577 \times 35) + (0.2423 \times 37)$$
$$A_r(\text{Cl}) = 26.52 + 8.97 = 35.49 \approx 35.5$$

This matches the value you see on the periodic table. Note that A_r is never a whole number when two or more isotopes are present in significant amounts.

Common mistake: Using percentage instead of fractional abundance. Always divide percentages by 100 before substituting. Using 75.77 instead of 0.7577 gives an answer 100× too large.

Mass Spectrometry

A **mass spectrometer** separates particles by their **mass-to-charge ratio** (m/z). It is used to:

1. Identify the isotopes present in an element and their relative abundances
2. Determine the A_r of an element
3. Identify unknown compounds (beyond SL scope)

How a mass spectrometer works (five stages):

Stage	Process	What happens
1. Vaporisation	Sample is converted to a gas	Particles are separated so they can be manipulated individually
2. Ionisation	Particles bombarded with high-energy electrons	Electrons are knocked off atoms, forming positive ions (M^+)
3. Acceleration	Ions accelerated through an electric field	All ions gain the same kinetic energy; lighter ions travel faster
4. Deflection	Ions passed through a magnetic field	Ions are deflected; lighter/more charged ions deflect more
5. Detection	Ions strike a detector	A signal proportional to the number of ions is recorded at each m/z ratio

Reading a mass spectrum:

- The x-axis shows m/z (for singly charged ions, this equals the mass number)
- The y-axis shows relative abundance (%)
- Each peak represents one isotope
- The height of each peak is proportional to the natural abundance of that isotope

When peak heights (not percentages) are given, use this form to calculate A_r — it normalises the abundances automatically:

$$A_r = \frac{\sum(m/z \times \text{relative abundance})}{\sum \text{relative abundance}}$$

IB TIP

IB Tip: You are expected to be able to read a mass spectrum and calculate A_r from it. In the exam, relative abundances may be given as peak heights rather than percentages — the calculation method is identical.

3. Electron Configurations (S1.3)

Energy Levels and Sublevels

Electrons in an atom occupy **energy levels** (principal quantum levels, $n = 1, 2, 3, 4, \ldots$). Within each energy level there are **sublevels** labelled s, p, d, f.

Sublevel	Number of orbitals	Maximum electrons	Shape
s	1	2	Spherical
p	3	6	Dumbbell-shaped (3 orientations: p_x, p_y, p_z)
d	5	10	Complex (5 orientations)
f	7	14	Complex (7 orientations)

Each **orbital** can hold a maximum of **2 electrons** (with opposite spins — Pauli exclusion principle).

Sublevels present in each principal energy level:

n	Sublevels present	Total electrons
1	$1s$	2
2	$2s, 2p$	8
3	$3s, 3p, 3d$	18
4	$4s, 4p, 4d, 4f$	32

The Three Rules for Filling Orbitals

1. Aufbau principle ("building up"): Electrons fill orbitals in order of increasing energy, starting with the lowest-energy sublevel.

The order of filling is:

$$1s \rightarrow 2s \rightarrow 2p \rightarrow 3s \rightarrow 3p \rightarrow 4s \rightarrow 3d \rightarrow 4p \rightarrow 5s \rightarrow 4d \rightarrow 5p \ldots$$

Note that $4s$ fills **before** $3d$ because $4s$ is lower in energy than $3d$ for neutral atoms.

2. Hund's rule: When electrons occupy orbitals of the same energy (degenerate orbitals), each orbital receives one electron before any orbital receives a second. All singly occupied orbitals have the same spin direction.

For example, in the $2p$ sublevel of nitrogen ($2p^3$), each of the three p orbitals has one electron — none share an orbital yet.

3. Pauli exclusion principle: No two electrons in an atom can have the same set of four quantum numbers. In practice, this means each orbital holds a maximum of **two electrons with opposite spins** (represented as ↑↓).

Writing Electron Configurations

Full configuration: List every sublevel and its electron count as a superscript.

Examples:

Element	Z	Full electron configuration
Hydrogen	1	$1s^1$
Helium	2	$1s^2$
Lithium	3	$1s^2\,2s^1$
Carbon	6	$1s^2\,2s^2\,2p^2$
Nitrogen	7	$1s^2\,2s^2\,2p^3$
Oxygen	8	$1s^2\,2s^2\,2p^4$
Neon	10	$1s^2\,2s^2\,2p^6$
Sodium	11	$1s^2\,2s^2\,2p^6\,3s^1$
Magnesium	12	$1s^2\,2s^2\,2p^6\,3s^2$
Chlorine	17	$1s^2\,2s^2\,2p^6\,3s^2\,3p^5$
Argon	18	$1s^2\,2s^2\,2p^6\,3s^2\,3p^6$
Calcium	20	$1s^2\,2s^2\,2p^6\,3s^2\,3p^6\,4s^2$
Iron	26	$1s^2\,2s^2\,2p^6\,3s^2\,3p^6\,4s^2\,3d^6$

Condensed (noble gas core) notation: Replace the inner electrons with the symbol of the preceding noble gas in square brackets.

Examples:

- Sodium: $[\mathrm{Ne}]\,3s^1$
- Sulfur: $[\mathrm{Ne}]\,3s^2\,3p^4$
- Calcium: $[\mathrm{Ar}]\,4s^2$
- Iron: $[\mathrm{Ar}]\,4s^2\,3d^6$

WORKED EXAMPLE

Worked Example 2 — Writing electron configurations

Write the full and condensed electron configurations for phosphorus ($Z = 15$).

Step 1: Total electrons = 15 (neutral atom).

Step 2: Fill in order: $1s^2\,2s^2\,2p^6\,3s^2\,3p^3$

Check: $2 + 2 + 6 + 2 + 3 = 15$ ✓

Step 3: Condensed — the noble gas before P is neon ($Z = 10$, configuration $1s^2\,2s^2\,2p^6$):

$[\mathrm{Ne}]\,3s^2\,3p^3$

Applying Hund's rule to the $3p^3$: The three $3p$ electrons each occupy a separate orbital with parallel spins before any pairing occurs:

$$3p: \uparrow \quad \uparrow \quad \uparrow$$

Exceptions: Chromium and Copper

Two important exceptions to the Aufbau filling order arise from the extra stability of **half-filled** (d^5) and **fully-filled** (d^{10}) d subshells.

Element	Expected	Actual	Reason
Chromium ($Z = 24$)	$[\mathrm{Ar}]\,4s^2\,3d^4$	$[\mathrm{Ar}]\,4s^1\,3d^5$	Half-filled d^5 is extra stable
Copper ($Z = 29$)	$[\mathrm{Ar}]\,4s^2\,3d^9$	$[\mathrm{Ar}]\,4s^1\,3d^{10}$	Fully-filled d^{10} is extra stable

⚠ EXAM ALERT

Exam trap — Cr and Cu exceptions: These are the two most commonly tested exceptions in IB Chemistry. If an exam question asks for the electron configuration of Cr or Cu, you must give the actual configuration, not the predicted one. For all other transition metals (Ti, V, Mn, Fe, Co, Ni, Zn), the regular Aufbau rule applies without exception.

Relating Electron Configuration to the Periodic Table

The periodic table is organised to reflect electron configurations:

- **s-block** (Groups 1–2): outermost electrons fill an s sublevel
- **p-block** (Groups 13–18): outermost electrons fill a p sublevel
- **d-block** (Groups 3–12, transition metals): outermost electrons fill a d sublevel
- **f-block** (lanthanides and actinides): outermost electrons fill an f sublevel

The **period number** gives the highest principal quantum level (n) occupied by electrons in the ground state. The **group number** (for main-group elements) gives the number of valence electrons.

MEMORISE THIS

Shortcut for valence electrons:

- Group 1: 1 valence electron (ns^1)
- Group 2: 2 valence electrons (ns^2)
- Groups 13–18: (group number − 10) valence electrons — e.g. Group 17 has 7 valence electrons ($ns^2\,np^5$)
- Noble gases (Group 18): 8 valence electrons (except He: 2)

The number of valence electrons determines bonding behaviour — this is why elements in the same group have similar chemistry.

Emission Spectra as Evidence for Discrete Energy Levels

When atoms are excited (by heat or electricity), electrons jump to higher energy levels. When they fall back to lower levels, they emit photons of light with specific **frequencies**. This produces a **line emission spectrum** — a series of discrete (separate) coloured lines on a dark background.

Key points:

- The lines are **discrete** (not continuous), proving that electrons can only occupy specific energy levels.
- Each element has a **unique** line spectrum — a "fingerprint" that can be used for identification (e.g. spectroscopy of stars).
- The energy of the emitted photon equals the energy difference between the two levels: $\Delta E = hf$ (where h is Planck's constant and f is frequency).
- The visible lines in hydrogen's spectrum (Balmer series) arise from transitions to the $n = 2$ level from $n = 3, 4, 5, 6$.

> 💡 **IB TIP**
>
> **IB Tip:** You do not need to calculate $\Delta E = hf$ at SL, but you must be able to explain *why* line spectra provide evidence for discrete energy levels. The key word is "discrete" — a continuous spectrum would be produced if electrons could have any energy.

4. Counting Particles by Mass — The Mole (S1.4)

Atoms are too small to count individually. Chemists use the **mole** as a counting unit — one mole of any substance contains the same number of particles as there are atoms in exactly 12 g of carbon-12.

Avogadro's Constant

$$N_A = 6.022 \times 10^{23}\ \text{mol}^{-1}$$

One mole of any substance contains 6.022×10^{23} particles (atoms, molecules, ions, etc.).

Molar Mass

The **molar mass** (M) of a substance is the mass of one mole of that substance, expressed in g mol^{-1}. Numerically, it equals the A_r or M_r (relative formula mass) in grams.

- $M(\text{C}) = 12.01\,\text{g mol}^{-1}$
- $M(\text{H}_2\text{O}) = 2(1.01) + 16.00 = 18.02\,\text{g mol}^{-1}$
- $M(\text{NaCl}) = 22.99 + 35.45 = 58.44\,\text{g mol}^{-1}$

The Mole Equation

$$n = \frac{m}{M}$$

Where:

- n = amount of substance (mol)

- m = mass (g)
- M = molar mass (g mol^{-1})

Rearrangements: $m = n \times M$ and $M = \frac{m}{n}$

✎ WORKED EXAMPLE

Worked Example 3 — Mole calculations

Part (a): How many moles are in 9.80 g of sulfuric acid, H_2SO_4?

Step 1 — Find molar mass: $M(\text{H}_2\text{SO}_4) = 2(1.01) + 32.07 + 4(16.00) = 2.02 + 32.07 + 64.00 = 98.09\,\text{g mol}^{-1}$

Step 2 — Apply $n = m/M$: $n = \frac{9.80}{98.09} = 0.0999 \approx 0.100\,\text{mol}$

Part (b): What mass of sodium chloride (NaCl) contains 1.20×10^{24} formula units?

Step 1 — Find moles: $n = \frac{1.20\times10^{24}}{6.022\times10^{23}} = 1.99 \approx 2.00\,\text{mol}$

Step 2 — Find mass: $m = n \times M = 2.00 \times 58.44 = 116.9 \approx 117\,\text{g}$

Relative Formula Mass

- **Relative formula mass** (M_r): sum of A_r values for all atoms in the formula unit — used for ionic compounds and molecules alike (e.g. $M_r(\text{H}_2\text{O}) = 2(1.01) + 16.00 = 18.02$)
- Both A_r and M_r are **dimensionless** (they are ratios relative to $\frac{1}{12}$ of the mass of ^{12}C, not absolute masses)

Empirical and Molecular Formulae

The **empirical formula** gives the simplest whole-number ratio of atoms of each element in a compound. The **molecular formula** gives the actual number of atoms of each element in one molecule.

- Molecular formula is always a whole-number multiple of the empirical formula.
- The multiple is found from: $\text{multiple} = \frac{M_r(\text{molecular})}{M_r(\text{empirical})}$

WORKED EXAMPLE

Worked Example 4 — Empirical and molecular formula from percentage composition

A compound has the following percentage composition by mass: C 40.00%, H 6.72%, O 53.28%. Its molar mass is $M_r = 90.08\,\text{g mol}^{-1}$. Find its empirical and molecular formulae.

Step 1: Assume 100 g of compound — percentages become masses in grams:

- $m(\text{C}) = 40.00\,\text{g}, m(\text{H}) = 6.72\,\text{g}, m(\text{O}) = 53.28\,\text{g}$

Step 2: Convert to moles: $n(\text{C}) = \frac{40.00}{12.01} = 3.331 \quad n(\text{H}) = \frac{6.72}{1.008} = 6.667 \quad n(\text{O}) = \frac{53.28}{16.00} = 3.330$

Step 3: Divide by the smallest value (3.330): $\text{C}: \frac{3.331}{3.330} \approx 1 \quad \text{H}: \frac{6.667}{3.330} \approx 2 \quad \text{O}: \frac{3.330}{3.330} = 1$

Empirical formula: CH_2O — $M_r(\text{empirical}) = 12 + 2 + 16 = 30\,\text{g mol}^{-1}$

Step 4: Find the multiple: $\text{multiple} = \frac{90.08}{30} = 3.003 \approx 3$

Molecular formula: $\text{C}_3\text{H}_6\text{O}_3$

EXAM ALERT

Exam trap — empirical formula ratios: After dividing by the smallest mole value, you may get a ratio like 1 : 1.5 : 2. Do not round 1.5 to 2 — multiply all values by 2 to get whole numbers (1 : 1.5 : 2 becomes 2 : 3 : 4). Only round to the nearest integer if the value is within 0.1 of a whole number.

5. Ideal Gases (S1.5)

An **ideal gas** is a hypothetical gas whose molecules:

- Have negligible volume (treated as point masses)
- Have no intermolecular forces between them
- Collide elastically (kinetic energy is conserved)

Real gases approximate ideal behaviour at **low pressure** and **high temperature**.

The Ideal Gas Law

$$PV = nRT$$

Where:

- P = pressure (Pa or kPa — be consistent with R)
- V = volume (m^3 or L — be consistent with R)

- n = amount of substance (mol)
- R = gas constant $= 8.314\,\text{J mol}^{-1}\text{K}^{-1}$ (use with Pa and m^3)
- T = temperature (K) — always convert: $T(\text{K}) = T(^\circ\text{C}) + 273.15$

Molar Volume at STP

Standard Temperature and Pressure (STP): $T = 273\,\text{K}$ (0°C), $P = 100\,\text{kPa}$

At STP, one mole of any ideal gas occupies a volume of **22.7 L** ($22.7 \times 10^{-3}\,\text{m}^3$).

$$V_m = \frac{RT}{P} = \frac{8.314 \times 273}{100\,000} = 0.02270\,\text{m}^3 = 22.7\,\text{L mol}^{-1}$$

⚠ EXAM ALERT

Exam trap — STP definition: The IB (from 2016 onwards) defines STP as 273 K and **100 kPa** (not 101.3 kPa). This gives $V_m = 22.7\,\text{L mol}^{-1}$, **not** $22.4\,\text{L mol}^{-1}$ (which was the old definition at 101.3 kPa). Using 22.4 in an IB exam will cost you marks.

WORKED EXAMPLE

Worked Example 5 — Ideal gas calculation

Calculate the volume occupied by 3.20 g of oxygen gas (O_2) at 298 K and 101 kPa. Give your answer in litres.

Step 1: Find moles: $M(\text{O}_2) = 2 \times 16.00 = 32.00\,\text{g mol}^{-1}$

$n = \frac{3.20}{32.00} = 0.100\,\text{mol}$

Step 2: Apply $PV = nRT$. Use $R = 8.314\,\text{J mol}^{-1}\text{K}^{-1}$, P in Pa:

$V = \frac{nRT}{P} = \frac{0.100 \times 8.314 \times 298}{101\,000} = \frac{247.8}{101\,000} = 2.45 \times 10^{-3}\,\text{m}^3$

Step 3: Convert to litres ($1\,\text{m}^3 = 1000\,\text{L}$):

$V = 2.45 \times 10^{-3} \times 1000 = 2.45\,\text{L}$

6. Periodic Trends — Atomic Radius, Ionisation Energy & Electronegativity

Although the IB 2025 syllabus places periodic trends formally in Structure 3, they are commonly examined alongside Structure 1 content and are essential for understanding bonding (Structure 2). This section covers the three most important trends.

Atomic Radius

The **atomic radius** is defined as half the distance between the nuclei of two identical atoms in a covalent bond (covalent radius).

Trend across a period (left → right): decreases

As you move across a period, the nuclear charge Z increases (more protons added), but the electrons are added to the **same energy level** (same principal quantum number n). The increased nuclear charge pulls all the electrons closer to the nucleus, so the atomic radius decreases.

Trend down a group (top → bottom): increases

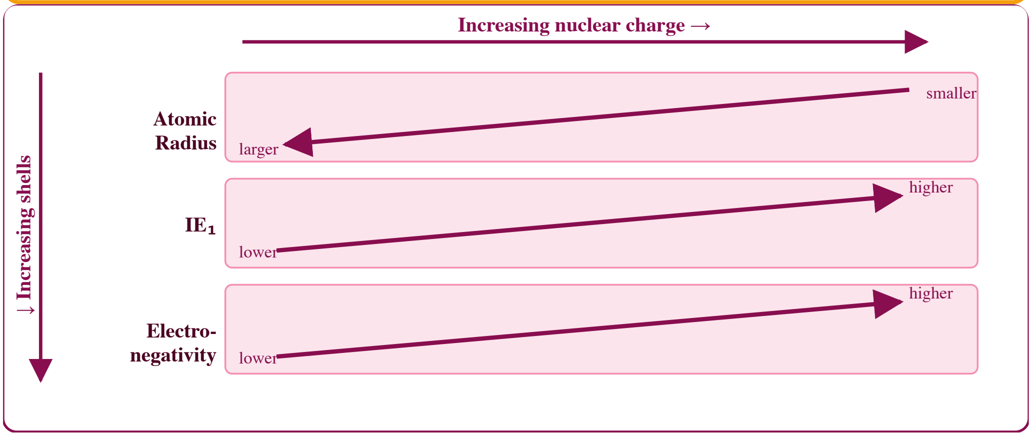
As you move down a group, each successive element has electrons in an **additional energy level** (higher n). The outermost electrons are further from the nucleus, so the atomic radius increases. Additionally, electrons in inner shells **shield** the outer electrons from the full nuclear charge (shielding effect).

MEMORISE THIS

Periodic trends summary table:

Property	Across period (→)	Down group (↓)	Reason
Atomic radius	Decreases	Increases	Nuclear charge vs. new energy levels
1st ionisation energy	Generally increases*	Decreases	Effective nuclear charge vs. shielding
Electronegativity	Increases	Decreases	Nuclear attraction on bonding electrons

*With exceptions: see Section 6.2



Periodic trends: atomic radius, ionisation energy and electronegativity change directionally across periods and down groups.

First Ionisation Energy (IE_1)

The **first ionisation energy** is the energy required to remove one mole of electrons from one mole of gaseous atoms in the ground state:

$$\mathrm{X}(g) \rightarrow \mathrm{X}^{+}(g) + e^{-} \quad \Delta H = \mathrm{IE}_1$$

IE_1 is always **endothermic** (energy must be supplied to overcome the attraction between the electron and nucleus).

Trend across a period: general increase

As nuclear charge increases across a period (electrons in the same shell), the effective nuclear charge experienced by the outer electrons increases. The outer electrons are held more tightly, so more energy is needed to remove them.

Anomalies across Period 2 and Period 3:

There are two important dips in the general increasing trend:

Anomaly 1 — Group 2 to Group 13 (e.g. Be → B, Mg → Al):

$IE_1(B) < IE_1(Be)$, even though B has higher Z.

- Beryllium: outer electron is in $2s$ ($1s^2\,2s^2$)
- Boron: outer electron is in $2p$ ($1s^2\,2s^2\,2p^1$)
- The $2p$ sublevel is **higher in energy** and **further from the nucleus** than $2s$, and is also **partially shielded** by the filled $2s$ electrons. Therefore the $2p$ electron in B is easier to remove than the $2s$ electron in Be.

Anomaly 2 — Group 15 to Group 16 (e.g. N → O, P → S):

$IE_1(O) < IE_1(N)$, even though O has higher Z.

- Nitrogen: $2p^3$ — each $2p$ orbital has one electron (Hund's rule)
- Oxygen: $2p^4$ — one $2p$ orbital must hold **two electrons** (paired)
- Electron–electron repulsion between the paired electrons in oxygen makes it **easier** to remove one electron from O than from N.

⚠ **EXAM ALERT**

Exam trap — IE anomalies: These two anomalies (B < Be and O < N) are among the most frequently tested points in IB Chemistry. You must be able to explain them using sublevel energies and electron pairing repulsion. Simply saying "because of the electron configuration" earns no marks — you must explain the specific reason (higher sublevel / electron repulsion from pairing).

Trend down a group: decreases

Down a group, the outer electron is in successively higher energy levels (greater n), further from the nucleus. Additionally, there are more inner electron shells providing greater **shielding** from the nuclear charge. Both factors reduce the effective nuclear charge felt by the outer electron, so less energy is needed to remove it.

Electronegativity

Electronegativity is a measure of the ability of an atom in a covalent bond to attract the shared pair of electrons towards itself. It is measured on the **Pauling scale**, where fluorine (the most electronegative element) is assigned a value of 4.0.

Trend across a period: increases

Moving across a period, nuclear charge increases while the atomic radius decreases (electrons are closer to the nucleus). The nucleus exerts a stronger attraction on the bonding electrons, so electronegativity increases.

Trend down a group: decreases

Moving down a group, the atomic radius increases and the shielding effect increases. The nucleus has less effective attraction over the bonding electrons, so electronegativity decreases.

> **IB TIP**
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> **IB Tip:** Noble gases are not typically assigned electronegativity values because they do not generally form covalent bonds. Caesium (Cs) has the lowest electronegativity of commonly discussed elements; fluorine (F) has the highest. Electronegativity differences are used to determine bond polarity and bond type — you will use this concept extensively in Structure 2 (Bonding & Structure).

Practice Questions

Test your understanding with these IB-style multiple choice questions.

Question 1

Rutherford's gold-foil experiment provided evidence for which of the following?

A. Electrons are present in discrete energy levels.

B. The atom contains a small, dense, positively charged nucleus.

C. Atoms are mostly empty space with electrons embedded in positive charge.

D. Neutrons and protons have approximately equal masses.

Question 2

Which of the following is the correct electron configuration for chromium ($Z = 24$)?

A. $[\mathrm{Ar}]\ 4s^2\ 3d^4$

B. $[\mathrm{Ar}]\ 4s^2\ 3d^2$

C. $[\mathrm{Ar}]\ 4s^1\ 3d^5$

D. $[\mathrm{Ar}]\ 4s^0\ 3d^6$

Question 3

An element has two isotopes: ^{63}X with 69.2% natural abundance and ^{65}X with 30.8% natural abundance. What is the relative atomic mass of this element?

A. 63.5

B. 64.0

C. 63.6

D. 64.4

Question 4

Which of the following correctly explains why the first ionisation energy of oxygen is lower than that of nitrogen?

A. Oxygen has a higher nuclear charge than nitrogen.

B. Oxygen has a larger atomic radius than nitrogen.

C. The paired electron in oxygen's $2p$ sublevel experiences greater electron–electron repulsion, making it easier to remove.

D. Oxygen's $2p$ sublevel is at a lower energy than nitrogen's.

Question 5

A sample of gas occupies $4.50 \times 10^{-3}\ \mathrm{m}^3$ at a temperature of 300 K and a pressure of $1.20 \times 10^5\ \mathrm{Pa}$. What is the amount of substance (in mol) present in the sample?

A. 0.180 mol

B. 0.216 mol

C. 0.270 mol

D. 1.80 mol

▶ Show Answers

Next Topic

The concepts in this guide — especially electron configurations and electronegativity — are directly applied in the next topic:

Bonding & Structure (Structure 2) — ionic bonding, covalent bonding, VSEPR geometry, intermolecular forces, and the bonding triangle all build directly on electron configurations and periodic trends covered here.